1. ELECTROCHEMISTRY - GALVANIC CELLS

S U M M E R  2016

Electrochemistry is considered to be a difficult topic in introductory chemistry. Even students who are able to successfully do quantitative problems on this topic often show misconceptions when asked to qualitatively describe the processes that occur in a galvanic or electrolytic cell. Yet it is important. Development of electrochemical technology for rechargeable batteries has led to the success of wireless technology. Electrochemistry will be instrumental in reducing our dependence on limited oil supplies through the development of fuel cells for cars and storage cells for solar power. Understanding electrochemistry also enables the prediction of the spontaneous direction of oxidation-reduction (redox) reactions. These reactions are among the most important in both chemistry and biochemistry.

A brief review of this topic is presented here, but you will also need to refer to your textbook (Atkins, Jones, and Laverman, Chemical Principles, 6th ed., Ch. 14) or any other freshman chemistry textbook. In addition, there are good animations of galvanic cells on the internet.

In the prelab homework for this experiment, you will calculate the amounts of the salts you will need to make the solutions for the two half-cells in Part A. Record this calculation in your notebook to make it quicker for you to get started in the lab. Check the results with your lab partner and with your TA before actually making the solutions. You will also do these calculations before you come to lab and record them in your notebook.

Blackboard resources
- Experiment → Exp 1
  - Laboratory instructions. You are required to print these and bring them with you to lab.
  - Background information with link to an animation
  - Online video on concepts and techniques
  - Answers to the inlab worksheet are posted after lab
  - Link to a techniques video on using a volumetric flask to make solution with accurate concentrations from a solid. (You will need to do this on the lab skills exercise later this semester.)
- Practice problems that will help you with the post-lab assignment.

Similar information is available for all experiments. In addition, there is information on Blackboard that will be useful for all experiments. This includes
- A schedule of all experiments and tests
- Laboratory notebook instructions
- Information on how to use Chem21labs
- Information on spreadsheets and graphing
- Information on accuracy of equipment and significant figures.

Objectives
- Use a volumetric flask to make solutions with accurate concentrations from a solid. (You will need to do this on the lab skills exercise later this semester.)
- Construct a galvanic cell and identify its parts. This includes recognizing
  - direction of electron flow
  - chemical half-reactions at the electrodes
  - the identity of the anode and cathode
  - the contents of the beakers
- Use the specialized notation for describing galvanic cells.
- Predict the voltage for galvanic cells using standard reduction potentials.
- Use the Nernst equation to calculate the concentration dependence of the potential.
- Construct a galvanic cell that produces a potential due to concentration differences in the half-cells and use this to review solubility product concepts.

Galvanic Cells
In a galvanic cell, electrical potential is generated by the difference in electrical reduction potential of two half-cells so that electrons flow through the external wire from the anode where they are produced by the oxidation to the cathode, where they are used in the reduction. Batteries are galvanic cells. Galvanic and electrolytic cells have the same fundamental definition of anode and cathode and the same processes occur at the electrodes. In galvanic cells the reactions are spontaneous, caused by the different reduction potentials of the ions in the different half-cells. In electrolytic cells, as you will see in the next experiment, the reactions require the work supplied by an external voltage.

In this experiment, you will make several galvanic cells and measure the voltage produced. In
Part A, you will make one galvanic cell from copper and lead. In Part B, you will be given several electrodes and unknown salt solutions and you will be asked to identify the unknowns from the voltage produced by the half-cells when combined with the copper half-cell from Part A. In Part C you will make a concentration cell and use the voltage produced to determine the concentration of copper ions in a saturated solution of copper oxalate. This will be used to determine the K_{sp} of copper oxalate.

All oxidation-reduction reactions can be written as the sum of two half-reactions. One half-reaction involves oxidation, the loss of electrons. The other half-reaction involves reduction, the gain of electrons. In order to compare half-reactions, it is conventional to write them all as reduction reactions. They are then tabulated in lists that indicate the relative tendency for each half-reaction to occur. The higher the reduction potential, the more likely the reduction. The values that are assigned are only relative values since all that can be observed is the voltage of a complete reaction. The values are all referenced to the reduction of H^+ ions, which is arbitrarily given the potential of zero.

The following table of standard reductions potentials used in this experiment is taken from the Chem21labs reference pages.

### Standard Reduction Potentials at 25°C

<table>
<thead>
<tr>
<th>Half-Reaction</th>
<th>E^o(volts)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag^+ + e^- → Ag(s)</td>
<td>0.799</td>
</tr>
<tr>
<td>Cu^{2+} + 2e^- → Cu(s)</td>
<td>0.337</td>
</tr>
<tr>
<td>Pb^{2+} + 2e^- → Pb(s)</td>
<td>-0.126</td>
</tr>
<tr>
<td>Sn^{2+} + 2e^- → Sn(s)</td>
<td>-0.136</td>
</tr>
<tr>
<td>Zn^{2+} + 2e^- → Zn(s)</td>
<td>-0.763</td>
</tr>
<tr>
<td>Al^{3+} + 3e^- → Al(s)</td>
<td>-1.660</td>
</tr>
<tr>
<td>Mg^{2+} + 2e^- → Mg(s)</td>
<td>-2.370</td>
</tr>
</tbody>
</table>

Consider the apparatus shown here:

![Galvanic cell diagram](image)

Here M is the metal electrode in contact with a solution containing M^{m+} cations and X^- anions and N is a different metal electrode in contact with a different solution containing N^{n+} cations and Y^- anions. The lower case n and m denote the oxidation number of the metal. A salt bridge allows ions to pass into the cells to maintain electrical neutrality but prevents gross mixing of the two solutions. A voltmeter connected between the electrodes will show that a voltage difference exists between the two sides. The voltmeter reading will be positive when the positive terminal is attached to the electrode with the higher reduction potential, labeled M in this diagram. The other electrode is attached to the COM terminal on the voltmeter. Electrons flow through the voltmeter from the N electrode to the M electrode. The reaction at the N electrode is N(s) → N^{n+}(aq) + ne^- . The N electrode is the anode, because oxidation occurs at this site (the N loses electrons, becoming N^{n+}). The reaction at the M electrode is M^{m+}(aq) + me^- → M(s). The M electrode is the cathode since the M^{m+} ions are reduced at this site and gain electrons. In your experiment, you will be able to observe which electrode is the anode and which is the cathode by observing the connections to the terminals on the voltmeter that produce a positive reading.

Reactions at the anode and cathode take place simultaneously in the apparatus. At the same time, ions from the salt bridge travel to the solutions, keeping the solutions neutral. The electrons produced by the oxidation of N generate a current in the wire that travels through the voltmeter to the M electrode, where electrons are used in the reduction of the M^{m+} in the solution at the cathode. The voltmeter measures the difference in the potentials between the two half-reactions. A small lamp could be substituted for the voltmeter and the movement of electrons would cause the lamp to glow.

A major source of confusion is about what processes occur in galvanic (and electrolytic) cells. Here it helps to look at a picture and you should look at the pictures in sections 14.4 and 14.5 in your textbook or at the animation on the internet referenced on Blackboard. The important things to observe are that electrons flow in the external wire while ions, not electrons, travel in the solutions and in the salt bridge. The salt bridge is used to complete the circuit and maintain electrical neutrality in the solutions. When positive ions are created at the anode, negative ions will travel from the salt bridge into the solution on the anode side so that the solution remains electrically neutral overall. At the same time, the positively charged ions near the cathode are being reduced so positive ions from the salt bridge must travel into that solution so that it remains electrically neutral and the salt bridge remains electrically neutral as well.

The voltmeter cannot be used to measure the potential of a half-reaction by itself; it measures the
difference in potentials between two half-reactions. Consequently, you can only determine relative values for the half-reaction potentials. They are given absolute numerical values by reference to the hydrogen half-reaction:

\[ 2\text{H}^+ (\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) \]

for which the half-cell potential is arbitrarily set at zero. The standard cell in electrochemistry is one in which the half-cell is combined with a hydrogen electrode under standard conditions (concentrations = 1M). For example, if a standard copper half-cell is connected to a hydrogen half-cell, a potential difference of 0.337V is observed at 25°C. If a standard lead cell is connected to a hydrogen half-cell, a potential difference of -0.126 V is observed at 25°C.

The standard voltage of a galvanic cell composed of two different half-cells is calculated from the difference between the standard potentials of the two half-cells. For example, for the cell shown in the diagram on the previous page:

\[ \Delta E^\circ = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = E^\circ(M^{m+} | M) - E^\circ(N^{n+} | N) \]  \hspace{1cm} (1)

Electrochemistry uses specialized notation. Half-cells are denoted by a single vertical line denoting a metal-solution interface. A galvanic cell (composed of two half-cells) is written as:

\[ \text{N} | \text{N}^{n+}(\text{conc}) || \text{M}^{m+}(\text{conc}) | \text{M} \] \hspace{1cm} (2)

The double vertical lines between the half-cells denote the salt bridge or ion-permeable membrane. The anode is on the left and the cathode is on the right. Because the reduction happens at the cathode, in galvanic cells it is always the half-cell with the highest reduction potential. For an actual cell, the concentrations of the solutions are included since that affects the voltage, as will be discussed later.

The voltage of a battery is determined by the difference in the half-cell potentials. The voltage lasts until one of the reactions producing the difference in half-cell potentials has exhausted the reactants, at which point we say the battery is "dead".

**CONCENTRATION DEPENDENCE**

In this experiment, the concentrations of the solutions are not 1M, which makes the experiment performed under nonstandard conditions. The dependence of the voltage on concentration is given by the *Nernst equation*. A good discussion of concentration effects is given in section 14.9 of your text. This equation is based on thermodynamic considerations.

The Nernst equation for a full galvanic cell is

\[ \Delta E = \Delta E^\circ - \frac{RT}{nF} \ln Q \] \hspace{1cm} (3)

where \( \Delta E \) is the voltage of the cell, \( \Delta E^\circ \) is the voltage of the cell under standard conditions, \( n \) is the number of electrons transferred in the reaction, \( R \) is the universal gas constant 8.314472 J mol\(^{-1}\) m\(^{-1}\), \( T \) is the temperature in Kelvin, \( F \) is the Faraday constant 96,485.3399 and \( Q \) is the reaction quotient for the reaction. If you substitute these values including the conversion factor for natural log to base ten log and determine \( \Delta E \) at 25°C, you get the more familiar form of this equation

\[ \Delta E = \Delta E^\circ - 0.05917 \frac{RT}{nF} \ln_{10} Q \]. \hspace{1cm} (4)

However, you should use the more general form shown in equation 3 since you will be able to use an experimental value for the temperature.

For half-reactions, the Nernst equation can be written as

\[ E = E^\circ + \frac{RT}{nF} \ln Q_h \] \hspace{1cm} (5)

where \( Q_h \) is the reaction quotient for the half-reaction.

For example, for the lead-copper cell, the overall net ionic reaction is

\[ \text{Cu}^{2+}(\text{aq}) + \text{Pb}(s) \rightarrow \text{Cu}(s) + \text{Pb}^{2+}(\text{aq}) \] \hspace{1cm} (6)

For this reaction, \( Q = [\text{Pb}^{2+}]/[\text{Cu}^{2+}] \) and \( n=2 \), so the Nernst equation can be written as

\[ \Delta E = \Delta E^\circ - \frac{RT}{2F} \ln (\frac{[\text{Pb}^{2+}]}{[\text{Cu}^{2+}]}) \] \hspace{1cm} (7)

\( \Delta E^\circ \) for the lead-copper cell is +0.463 V from the standard reduction potentials of copper and lead. As long as the concentration of \( \text{Cu}^{2+} \) and \( \text{Pb}^{2+} \) are equal, the log term is zero. However, if the concentrations are not equal, the concentrations must be included to determine the theoretical voltage. The concentration ratio enters in the argument of the log term. This means that a large concentration differential produces only a small voltage change. For example, if you were to measure the voltage of the cell with \([\text{Pb}^{2+}] = 2.00 \) M and \([\text{Cu}^{2+}] = 0.200 \) M at 25°C, the voltage would be 0.463 – (0.05917/2)log(10) = 0.433 V, only a 6.5% difference in the voltage for a factor of 10 difference in the concentrations in the two cells.

The situation with respect to the concentrations is more complicated when the cations have different charges. For example, if the cation being oxidized has an oxidation number of +2, the half reaction is

\[ \text{N}(s) \rightarrow \text{N}^{2+}(\text{aq}) + 2\text{e}^- \] \hspace{1cm} (8)

If the anion has an oxidation number of +1, then the half reaction for the cation being reduced is

\[ \text{M}^+(\text{aq}) + \text{e}^- \rightarrow \text{M}(s) \] \hspace{1cm} (9)

When these are combined into a full cell, the reaction equation is written as

\[ \text{N}(s) + 2\text{M}^+(\text{aq}) \rightarrow \text{N}^{2+}(\text{aq}) + 2\text{M}(s) \] \hspace{1cm} (10)
The reaction quotient $Q = \frac{[N^{2+}]}{[M^+]^2}$ \hspace{1cm} (11)

with the denominator term squared due to the stoichiometry.

The Nernst equation is written as

$$
\Delta E = \Delta E^0 - \frac{RT}{2F} \ln \left( \frac{[N^{2+}]}{[M^+]^2} \right) \hspace{1cm} (12)
$$

again with the denominator in the log term squared. If non-standard concentrations are used in either of the half-cells, the log term must be considered in the calculation for the theoretical value of the voltage. If the half-cell chemical equations involve different numbers of electrons, then you must pay attention to the stoichiometry when you write the Nernst equation since it affects the powers in the reaction quotient.

Voltages also develop between two half-cells having the same chemical species but different concentrations. This will be the case in Part C of this experiment where you will make a galvanic cell consisting of one half-cell with the concentration of cupric ion of 0.05M and one half-cell with the concentration of cupric ion that is several orders of magnitude lower. In the lower concentration half-cell, the concentration will be limited by the solubility product of an insoluble copper salt. The half-cell potential for each half-cell will be

$$
E = 0.337 - \frac{RT}{2F} \ln \left( \frac{1}{[\text{Cu}^{2+}]} \right) \hspace{1cm} (13)
$$

The half-cell with the higher concentration of Cu$^{2+}$ will have the larger reduction potential, so it will be the cathode. As a result, for this special case,

$$
\Delta E = - \frac{RT}{2F} \ln \left( \frac{[\text{Cu}^{2+}]_{\text{anode}}}{[\text{Cu}^{2+}]_{\text{cathode}}} \right) \hspace{1cm} (14)
$$

since $\Delta E^0 = 0$ because the standard reduction potentials cancel.

SAFETY

Using safe laboratory practices is important when using unknowns. Some of the metal salts are toxic. Lead nitrate, for example, is a poison. Others, such as aluminum chloride and silver nitrate, are corrosive. Silver nitrate will also turn your skin black after it is exposed to light. Always wear gloves and goggles. Keep the chemicals under the hood and work with your face as far from the chemicals as possible.

EXPERIMENTAL PROCEDURE

You will have one lab period to complete all three parts of this experiment. Obtain a reaction plate with 6 wells, a multimeter, 2 alligator clips, and a glass thermometer from the stockroom along with two copper metal strips and one lead metal strip. You will also need a 20-ml volumetric flask, a 100-ml volumetric flask, and a 1-ml pipette and pipettor. Your TA will give you short pieces of string to use as salt bridges in Parts A and B and a strip of filter paper to use as the salt bridge in Part C.

Part A: Construct a lead-copper galvanic cell.

You will need to prepare two solutions:

CuSO$_4$ ~0.05M 100 ml quantity

Pb(NO$_3$)$_2$ ~0.05M 20 ml quantity

Determine the amounts of chemical you will need to make these solutions. Your final concentration should be within 10% of these concentrations. The copper (II) sulfate salt is hydrated, with the molecular formula CuSO$_4$·5H$_2$O. The lead nitrate salt is not hydrated. Check your calculations with your lab partner and answer the first question on the in-lab worksheet before making the solutions.

Note: To avoid long lines at the balances, you should only get one salt at a time. Make up the first solution before getting the second salt. By then, the lines will be gone. Students who monopolize the balances before everyone has had a chance to get at least one of the compounds will lose points.

Use the following procedure to fill volumetric flasks when you need to dissolve a compound and accurately know the concentration. A video showing this technique is on Blackboard. You will be tested individually on making solutions as part of the lab skills exercise later this semester.

1) Dissolve the measured amount of compound in a beaker, using about 50% of the volume of the volumetric flask. So, for example, you should dissolve the copper(II) sulfate in about 50 ml of water in a beaker. Also rinse any solid remaining in the weighing pan into the beaker using your wash bottle. Then pour that into the volumetric flask.

2) Rinse any remains in the beaker into the volumetric flask. Use about half the remaining volume. So, for example, you should rinse the copper sulfate beaker with about 20-25 ml of water and pour it into the volumetric flask.

3) Repeat step 2, using less water.

4) Use your wash bottle to carefully fill the volumetric flask to a few mm below the line. Be careful – it fills very quickly once you reach the narrow neck of the flask.

5) Use your eyedropper to fill the flask so that the bottom of the meniscus is on the line.

6) Then, cover the flask with parafilm®. Hold your finger over the parafilm® and invert the flask several times to fully mix the solution.

Make the lead(II) nitrate solution in a similar manner, except you will only be making 20 ml using the 20 ml volumetric flask.
Make the lead-copper cell by filling one of the two center wells with copper(II) sulfate solution and the other center cell with lead nitrate solution.

Polish the copper and lead electrode strips with emery paper until they are shiny. Rinse with water. Bend them so that they will be able to stay in the appropriate well in the reaction plate. Then connect the red and black multimeter leads to the copper and lead strips using the alligator clips. (Depending on your voltmeter, either turn the dial to V==, or make sure the button is in the “out” position. You will be measuring a direct current, DC, voltage, not an alternating current, AC, voltage.) It is conventional, but not essential, for the red cable to be connected to the + connection on the multimeter and the black cable to the COM (or negative) connection. The electrons that will be produced at the anode will flow through the cable to the multimeter, through the other cable and then to the cathode, where they will be used to reduce the metal ions to solid metal. Make sure all your connections are tight. Do not just insert the voltmeter leads into the end of the alligator clip, you need to clip the lead directly to the electrode to obtain a tight connection, as shown below.

Turn on the multimeter to the voltage position, being sure that your multimeter is set to measure direct current. You will be able to record your voltage to the most number of significant digits if you set the multimeter dial to the lowest voltage range at which the voltage is on scale. If your multimeter reading is negative, switch the leads on the electrodes. The voltage should settle quickly if your connections are clean and tight. If not, remove the electrodes, clean, and repeat. Record the voltage and record the temperature in one of the wells. Make a sketch of the apparatus that produces a positive voltage in your notebook showing the voltage and which connection on the voltmeter is attached to which electrode. This will help you identify which electrode is the anode and which is the cathode. In addition, you should complete the sketch for part A on your in-lab worksheet. You also need a sketch in your notebook, but the one in your notebook only needs labels for the + and com connections from the voltmeter, the contents of the half-cells, the metals in the electrodes, the temperature, and the voltage.

Part B: Unknowns

You will be given a sheet with your unknown assignments at the start of the lab. Bring this sheet with you to obtain your assigned 2 unknown electrodes and 2 unknown salt solutions from your TA. Label the beakers or flasks and also the wells you use for the unknown solutions since all are clear and colorless and the experiment will not be successful if you mix them up. Fill one of the center wells with copper (II) sulfate solution. Construct galvanic cells using each of your unknowns in wells adjacent to the copper cell. Use new salt bridges for each full galvanic cell to avoid contaminating the solutions in the wells. Measure the temperature in one of the cells. Connect the voltmeter so that you obtain a positive voltage reading for each unknown when combined with your copper half-cell. After you have recorded the temperature and voltage, make a drawing in your notebook showing the connections and the voltage for each of the cells. Include the unknown number with each of your sketches in your notebook since you will need that for your post-lab assignment.

In addition to the sketches in your notebook, you also need to make and label the sketches on your in-lab worksheet.
The unknowns will be taken from the following list: silver electrode with AgNO₃ solution, aluminum electrode with AlCl₃ solution, magnesium electrode with MgSO₄ solution, tin electrode with SnCl₂ solution, or zinc electrode with ZnSO₄ solution. The concentrations of the solutions are such that the log term in the Nernst equation can be ignored when identifying the unknowns based on calculating their theoretical voltages when combined with your copper half-cell.

Before you discard the solutions in the wells, identify the 2 unknowns you were assigned and turn the completed form in to your TA. You and your lab partner can figure this out working together, or each may do it separately. It’s up to you. Do not discard your solutions until you both have made the identifications. You will lose points for every unknown for which you need to get a repeat solution.

Return the electrodes to your TA when you are finished. The solutions of the unknowns should all go in the large drums for disposal. Save your remaining copper sulfate solution for part C. Answer question 2 on your in-lab worksheet before continuing.

**Part C: Concentration cell**

Construct a Cu²⁺ concentration cell with a large concentration ratio. You will do this in two small beakers, rather than in the reaction well plate.

In order to get a very small cupric ion concentration on one side of this cell so that a measurable voltage is developed, you will use a saturated solution of solid copper oxalate in water. Copper oxalate is only slightly soluble; the solubility product determines the concentration of the copper in the solution. The voltage measurement will allow the solubility product of the salt to be calculated.

Get 50 ml of the 0.1M solution of oxalic acid from a carboy. Prepare the copper oxalate solution by adding 2 ml of your ~0.05M CuSO₄ solution to the 50 ml of 0.1M oxalic acid solution. This solution has a large excess of oxalate ion compared to cupric ion. Stir the solution gently for at least 5 minutes to allow the reaction to reach equilibrium. The copper oxalate precipitate (CuC₂O₄) will probably appear only as a slight cloudiness in the solution or may not be visible at all. Even if you can’t see the precipitate, this solution now contains Cu²⁺ ions at a concentration determined by the solubility product of copper oxalate. This will be the solution used in one half-cell and the ~0.05M CuSO₄ solution you made in part A will be the solution for the other half-cell. Prepare a salt bridge using a piece of filter paper dipped in KNO₃ to connect the two half-cells. Polish two copper electrodes. Put one in each beaker and measure the voltage developed by the cell and the temperature in one of the beakers. Just as in part A, you should connect the leads to the electrodes so as to obtain a positive reading on the multimeter. This will help you identify which half-cell has the polarity and which has the cathode. Make a sketch in your notebook. Answer question 3 on the in-lab worksheet and then complete the sketch on the worksheet.

**DATA ANALYSIS FOR Kₛₚ CALCULATION**

The solubility product of copper oxalate is

\[ K_{ₚ} = [Cu^{2+}] [C₂O₄^{2-}] \] (15)

As part of the Chem21 labs assignment, you will determine the solubility product by determining the concentrations of the Cu²⁺ and C₂O₄²⁻ ions.

To determine [Cu²⁺] in the copper oxalate cell, you will use the measured voltage of the cell and the known concentration of the Cu²⁺ from the solution you made in Part A in the Nernst equation as shown in equation 13.

You determine [C₂O₄²⁻] from the ionization constants of oxalic acid. Oxalic acid (H₂C₂O₄) is a weak diprotic acid whose first ionization constant at 25°C is 5.9 x 10⁻² and whose second ionization constant at 25°C is 6.4 x 10⁻⁵. To calculate the concentration of C₂O₄²⁻ in the solution, you make use of the fact that the second ionization constant is much smaller than its first. The dissociation of oxalic acid can be written as follows:

\[ H₂C₂O₄ = H^+ + HC₂O₄⁻ \quad K_{a1} = 5.9 \times 10^{-2} \] (16)

\[ HC₂O₄⁻ = H^+ + C₂O₄²⁻ \quad K_{a2} = 6.4 \times 10^{-5} \] (17)

To a good approximation, all the H⁺ and HC₂O₄⁻ come from the first step and therefore have the same concentration. Since

\[ K_{a2} = [H^+][C₂O₄^{2-}]/[HC₂O₄^{-}] \] (18)

and \[ [H^+] = [HC₂O₄^{-}] \], (19)

it follows that

\[ K_{a2} = [C₂O₄^{2-}] \] (20)

This is a surprisingly simple relationship that sets the concentration of the fully dissociated conjugate base equal to the second ionization constant. It is true for all diprotic acids as long as the second ionization constant is much smaller than the first ionization constant.

**EXCESS CHEMICAL DISPOSAL**

Heavy metal ions can cause environmental damage. They are toxic to many forms of marine and plant life. All the metal ion solutions from Parts A and B should be neutral so you do not need to add baking soda before disposing of them in the excess chemical drums. Your solution containing copper oxalate must be neutralized since the oxalate ion
hydrolyzes to oxalic acid. Dispose of all these solutions in the excess chemical drums. Any leftover oxalic acid (not containing metal ions) should also be neutralized and poured into the drums. Rinse your beakers and volumetric flasks out with a large amount of tap water. Rinse your electrodes before returning them. Give the unknown electrodes back to your TA and the copper and lead electrodes back to the stockroom with the rest of the equipment. Your used salt bridges can be thrown in the trash.

**UNKNOWN IDENTIFICATION**

Before you leave the lab, turn in your completed unknown assignments to your TA along with the electrodes.

**ASSIGNMENT**

Complete the Chem21labs assignment. Doing the practice problems on Blackboard will help with some of the calculations.